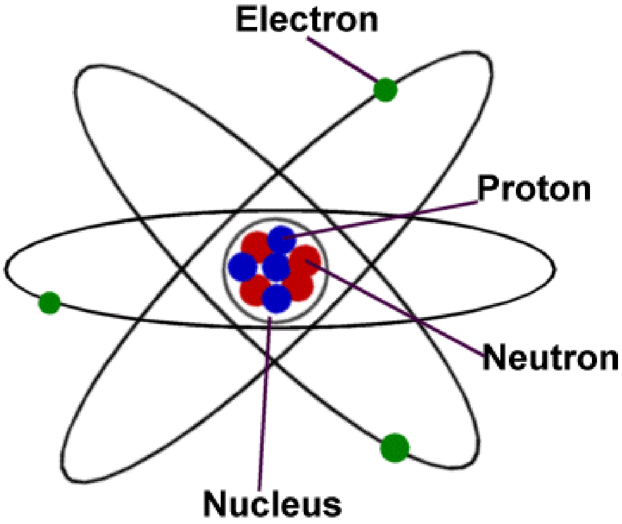
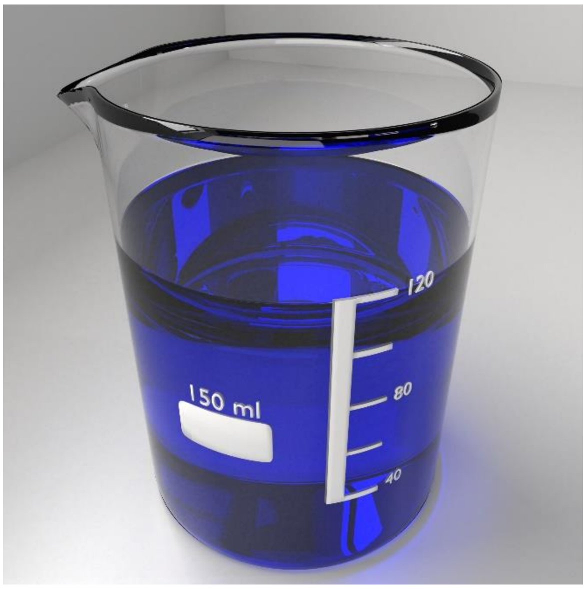


YEAR 11 CHEMISTRY – MODULE 1 NOTES

Properties and Structure of Matter

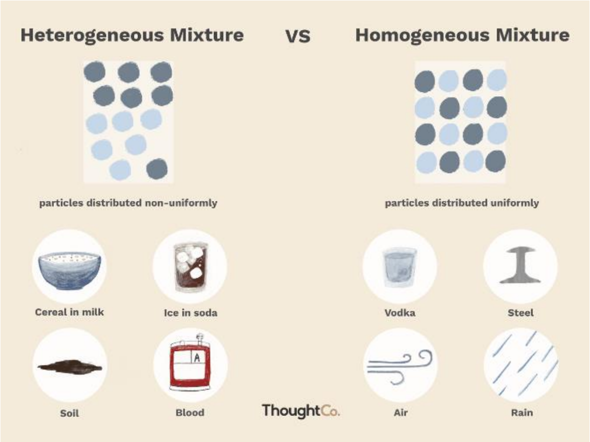


TOPIC 1: PROPERTIES OF MATTER

Homogenous and Heterogenous Mixtures

Homogenous Mixtures – *same* composition and properties throughout

Heterogenous Mixtures – *different* composition throughout



Physical Properties

Melting Point – Temperature reached to form liquid from solid

Boiling Point – Temperature reached to evaporate to gas from liquid

Density – Mass of atoms which make up substance, how close atoms are packed together

Solubility – Ability for substance to dissolve in water

Separation Techniques

Technique	*Property*
Filtration	A solid and liquid or solution
Evaporation	Liquid lower B. P than solid
Distillation	Big difference in B. P
Fractional Distillation	Small difference in B. P
Decantation	Density
Separating Funnel	Immiscible Liquids

Percentage Composition

$$\%_{\text{composition}} = \frac{\text{mass of component}}{\text{mass of compound}} \times 100$$

Naming Compounds

Inorganic Compounds – *doesn't* contain carbon

Elements – inorganic substances (same atomic number)

Compounds – two or more elements → chemically bonded

Ionic Compounds – metal and a non-metal

Covalent Compounds – 2 non-metals or non-metal/metalloid (semi metal)

Naming Covalent Compounds (2 non-metals)

- Name of element closest to left hand side written first
- Second part of name: use prefix based on atoms, then add -ide
- Prefixes (from 1 to 10 in order): mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca

e.g. Dinitrogen Tetroxide is N_2O_4

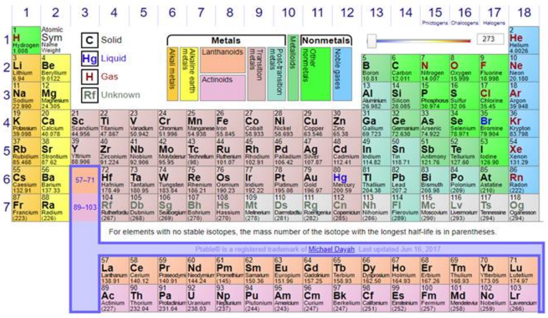
Naming Ionic Compounds

- Identify ions present – determine charge of each ion (refer to group for charge)
- Swap charges around (becomes the atoms of the other element/compound)
- Ends in -ide if monoatomic, or -ite/-ate if polyatomic (check polyatomic ion eg. SO_4^-)
- Roman numerals indicate charge of cation (when element is neutralized)

The Periodic Table

Periods – horizontal rows (7 of them)

Groups - (18 groups, transition metals are from 3-12)



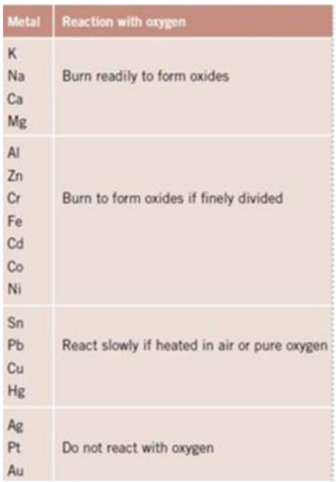
- Electrons in same group – similar physical and chemical properties

Physical Properties of Metals/Non-Metals

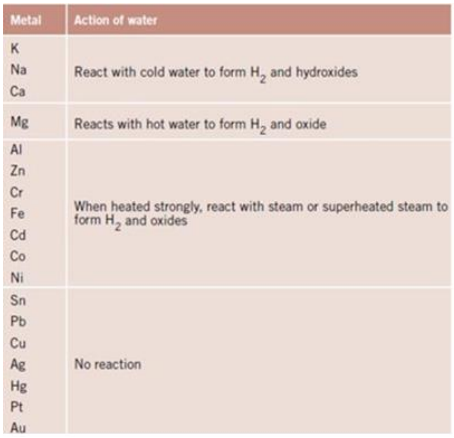
Metals	*Non-Metals*
• Shiny (lustrous) • Conductors of electricity/heat • Malleable (can be beaten or rolled into thin sheets) • Ductile (can be drawn out into wire by tensional forces) • Solids	• Can be solids, liquids, or gases at room temperature • Solids – dull in appearance • Poor electricity/ heat conductors • Brittle (smashes into small crystals) • Most exist as molecules rather than single atoms

Chemical Properties of Groups

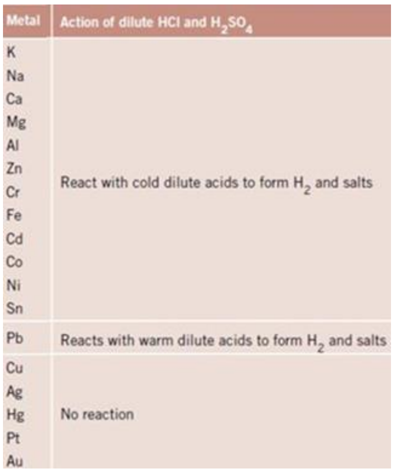
Group 1 (Alkali Metals)	Soft, highly reactive with water
Group 2 (Alkali Earth Metals)	Reacts with water
Group 3-12 (Transition Metals)	Less reactive than metals in Group 1 & 2
Group 17 (Halogens)	Combines with metals to form salts
Group 18 (Noble Gases)	Unreactive under normal conditions



Metal	Reaction with oxygen
K Na Ca Mg	Burn readily to form oxides
Al Zn Cr Fe Cd Co Ni	Burn to form oxides if finely divided
Sn Pb Cu Hg	React slowly if heated in air or pure oxygen
Ag Pt Au	Do not react with oxygen



Metal	Action of water
K Na Ca	React with cold water to form H_2 and hydroxides
Mg	Reacts with hot water to form H_2 and oxide
Al Zn Cr Fe Cd Co Ni	When heated strongly, react with steam or superheated steam to form H_2 and oxides
Sn Pb Cu Ag Hg Pt Au	No reaction



Metal	Action of dilute HCl and H_2SO_4
K Na Ca Mg Al Zn Cr Fe Cd Co Ni Sn	React with cold dilute acids to form H_2 and salts
Pb	Reacts with warm dilute acids to form H_2 and salts
Cu Ag Hg Pt Au	No reaction

TOPIC 2: ATOMIC STRUCTURE

Subatomic Particles

Protons – positively charged particles in the nucleus

Neutrons – neutral particles which are found in the nucleus, slightly heavier than protons

Electrons – negatively charged particles found outside nucleus. Exists in energy levels or shells surrounding the nucleus.

Isotopes

- Atomic number is written as subscript and nucleon number as a superscript before the symbol.

$$^{A}_{Z}E$$

A = number of protons + neutrons in nucleus

Z = number of protons in the nucleus

E = element

- Isotopes have same atomic number but different mass.

Electron Energy Levels and Orbits

- Electrons can move from lower shells to higher shells by absorbing energy equal to the energy between two shells
- When the electron is “excited” it moves to the next energy level
- When the electron is in its “emission” phase, emits the quantum of light and moves back to the previous shell.

Electron Configuration

- 1st shell can hold up to 2 electrons
- 2nd shell can hold up to 8 electrons
- 3rd shell can hold up to 18 electrons
- 4th shell can hold up to 32 electrons

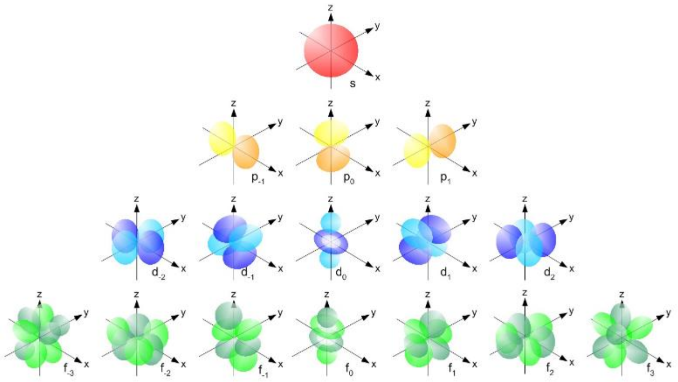
The valence shell is the outermost shell which contains electrons. They are stable if the number of electrons in the shell is 8 (an octet).

Metals form positively charged ions (cations) by loss of electrons.

Non metals form negatively charged ions (anions) by gain of electrons.

Subshells and Orbitals

- Electrons in each shell occupy subshells and orbitals.
- Orbitals are designated by the letters s, p, d, f.

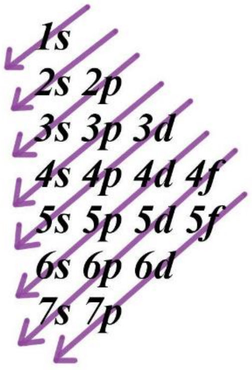


SPDF Notation

- The coefficient before the letter represents the energy level. E. g. 2*p* is on the n = 2 shell.
- The *s* orbital contains 2 electrons, *p* contains 6 electrons, *d* contains 10 electrons, *f* contains 14.

Electron configurations of elements can be expressed in this notation:

Hydrogen	$1s^1$
Lithium	$1s^2 2s^1$
Neon	$1s^2 2s^2 2p^6$
Sodium	$1s^2 2s^2 2p^6 3s^1$



Relative Atomic Mass

- The relative atomic mass is calculated by known isotope masses and the proportion of each isotope in the mixture

Example:

Naturally occurring chlorine consists of 75% chlorine-35 which has a relative atomic mass of 35.0, and 25% chlorine-37, which has a relative atomic mass of 37.0. Calculate the relative atomic mass of naturally occurring chlorine.

$$A_r(\mathrm{Cl}) = \frac{75A_r(\mathrm{Cl}-35) + 25A_r(\mathrm{Cl}-37)}{100}$$

$$A_r(\mathrm{Cl}) = \frac{75 \times 35.0 + 25 \times 37.0}{100}$$

$$A_r(\mathrm{Cl}) = 35.5$$

Flame Tests

Colours obtained from flame tests:

Element	*Flame Colour*
Copper	Blue
Lithium, Strontium	Red
Calcium	Orange
Sodium	Yellow
Barium	Green

- The energy difference from one energy level to another is represented by the colour it gives off.

Schrodinger and Bohr

Bohr proposed that electrons have discrete energies, and the atoms had discrete energy levels.

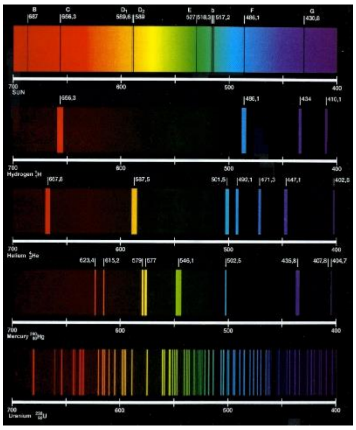
Due to his theory, he was able to get an idea of the spectrum of a hydrogen atom, but couldn't figure out more complex spectra.

Schrodinger used his equation. When he tried with more complex spectra, it led to the idea of sets of orbitals and sublevels.

Therefore, his equation was able to interpret emission spectra of atoms which have more electrons.

Emission Spectra

- Different energy light photons are seen as the coloured spectral lines in the emission spectrum.
- Greater the energy change, higher the frequency.
- All elements have different spectra because each element has different numbers of electrons in orbitals of varying energy



Nuclear Stability

- Protons and neutrons held in the nucleus by strong nuclear force
- Too many protons/neutrons for short ranged nuclear force – becomes unstable.

Radioisotopes

- Unstable isotopes emit radiation and are called radioisotopes. Emission of radiation continues until nucleus becomes stable

Alpha Particles: Composed of helium nuclei ($^{4}_{2}\mathrm{He}$)

Beta Particles: Electrons ($^{0}_{-1}\mathrm{e}$)

Gamma (γ) is electromagnetic radiation

- Stability of an isotope depends on neutron/proton ratio.

N/P ratio too high (excess neutrons):

- Forms proton and beta particle emitted from nucleus

$$^{131}_{53}\mathrm{I} \rightarrow {}^{123}_{54}\mathrm{Xe} + {}^{0}_{-1}\mathrm{e}$$

N/P ratio too low (excess protons):

- Forms proton and positron (electron with positive charge)

$$^{22}_{11}Na \rightarrow ^{22}_{10}Ne + ^{0}_{1}e$$

Nucleus too heavy:

- Alpha decay occurs – loss of an alpha particle reduces nucleus by 2 protons and 2 neutrons.

$$^{230}_{90}Th \rightarrow ^{226}_{88}Ra + ^{4}_{2}He$$

Production of Radioisotopes

- Nuclear Fission Reactors: reactors which nuclear fission reactions occur with production of energy. Bombards radioisotopes with neutrons

- Linear Accelerators: particle accelerators increasing energy of charged particles using electric fields. E.g. Og (Oganesson) was made by firing a beam of Calcium-48 ions at a californium-249 target.

- Cyclotrons are circular particle accelerators which increases energy of charged particles using static magnetic fields and electric fields.

- Half-life: The time for a radioisotope to decrease in mass by 50%.

TOPIC 3: PERIODICITY

States of Matter

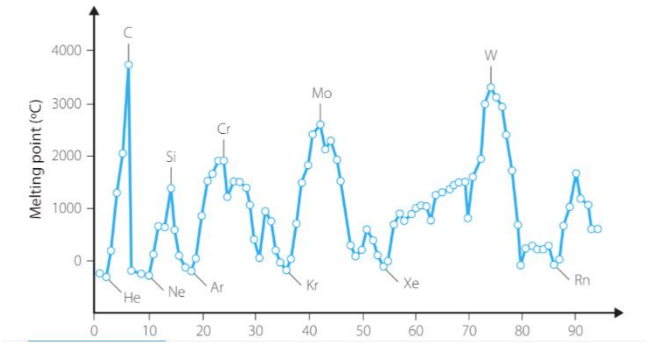
- *Metals are solids* at room temperature
- Non metals can be solid, liquid, or gas at room temperature
- Noble gases are all gases

EXCEPTIONS:

- Hydrogen is gas at room temperature
- Bromine and Mercury are liquid at room temperature

In general:

- Solid – if melting point greater than room temperature
- Liquid – if melting point lower but boiling point higher than room temperature
- Gas – if boiling point is lower than room temperature



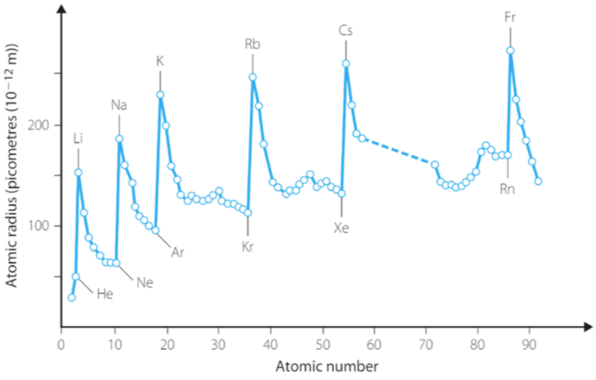
Electronic Configuration and Atomic Radii

- The group of an element determines the valency (how many electrons are in the outermost shell of an element)

- Atomic radius is the distance between the nucleus and the valence shell of an element

- *Atomic radius increases along a group* because there are more shells and hence the valence shell is further away from the nucleus

- *Atomic radius decreases along a period* because there is more attraction between the electrons and the protons – number of protons and electrons increase.



Electronegativity

- Electronegativity determines how likely an atom is to attract an electron.

Elements with fewer electrons find it *easier to lose an electron rather than gain*
Elements with many electrons find it *easier to gain new electrons rather than lose*

- The greater the atomic radii, less electronegative atom will be, nucleus and valence shell don't attract as much.

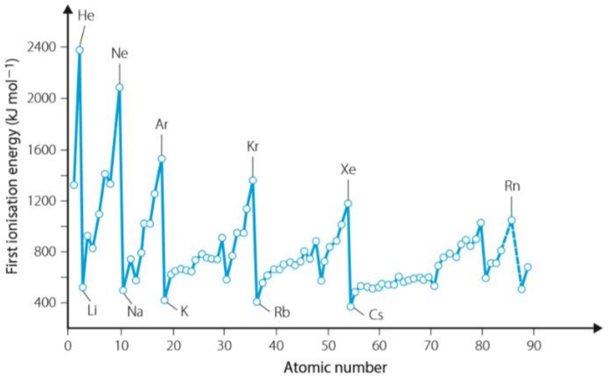
- Moving across a period increases electronegativity, moving down a group decreases electronegativity

First Ionisation Energy

- The ionisation energy is the energy required to rip an electron from the atom.

- Electrons less attracted to the nucleus can be ripped more easily than those which are more attracted

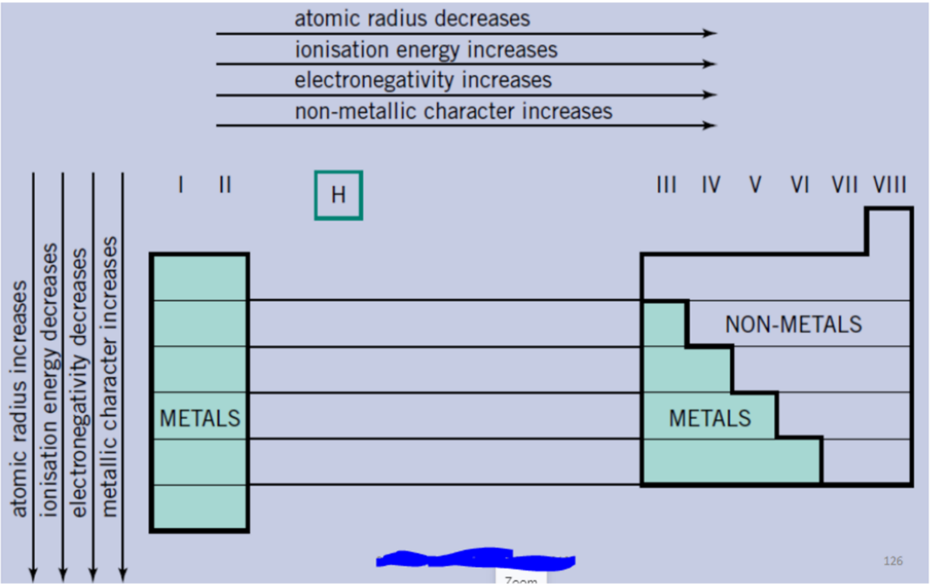
- Moving down a group decreases ionisation energy, moving across a period increases ionisation energy



Reactivity

- As you move down a metallic group, elements become more reactive
- As you move across, metallic character decreases

Summary of Periodicity



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TOPIC 4: BONDING

Ionic Bonding

- Occurs when an atom transfers its valence electrons to another atom. These bonds are strong.

Covalent Bonding

- Occurs when one atom shares its electrons with another atom, usually restricted to non-metals

Steps for Drawing Lewis Dot Structures

1. Count valence electrons in element/molecule
2. Determine the central atom (least electronegative usually – except hydrogen)
3. Draw single bonds to central atom (each single bond worth 2 electrons)
4. Put all remaining valence electrons on atoms as lone pairs. Electrons placed on most electronegative first
5. Turn lone pairs into double/triple bonds to give every atom an octet

Determining Formal Charge for LDS

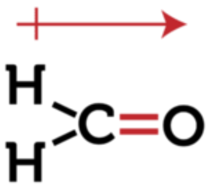
- Formal charge is used to determine the best Lewis Dot Diagrams for a molecule
- Make charges as close to zero as possible

$$FC = \text{valence electrons} - \text{non-bonding electrons} - \frac{\text{bonding electrons}}{2}$$

Polarity

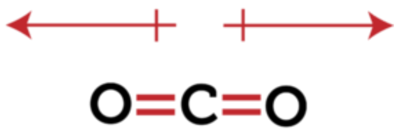
- Measure of charge distribution
- Polar substances have a positive end and a negative end
- Non-polar substances are neutral.

Formaldehyde (CH_2O)

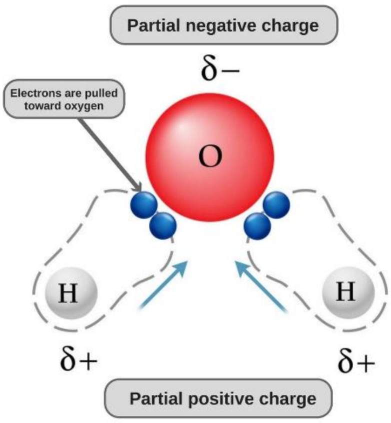


Formaldehyde is a polar compound. One end of the molecule has a slightly positive charge. The other end has a slightly negative charge.

Carbon Dioxide (CO_2)



Carbon dioxide is a nonpolar compound. Both ends of the molecule are slightly negative in charge.

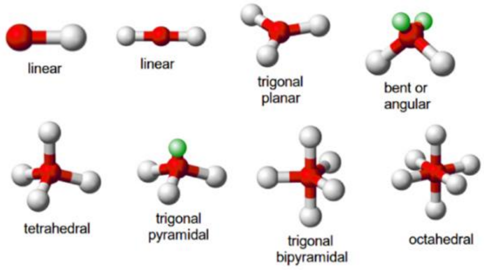


- Delta symbols show polarity of a substance.
- This is polar, has a separation of a positive and negative charge.

VSPER Theory (Valence Shell Electron Pairing Repulsion Theory)

- Pairs of valence electrons arrange themselves spatially
- Aim is to get as far away from one atom as possible

Different shapes:



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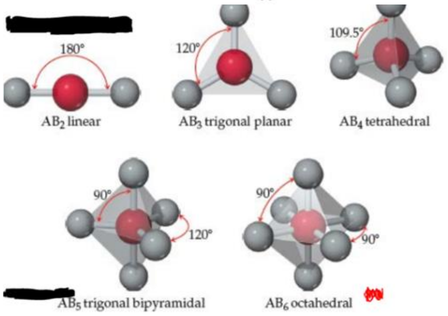

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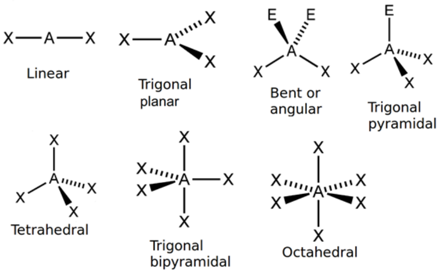
3D Structure:

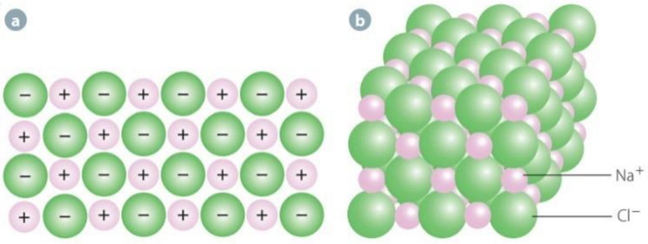


Comparison of Properties of Ionic and Covalent Molecular Substances

Ionic	*Covalent-molecular*
Always compounds	Exists as elements or compounds
Solid at room temp	Mostly gases or liquids, few are solid
High melting and boiling points	Low melting and boiling points
Hard and brittle	Does not conduct electricity as solids or liquids
Do not conduct electricity as solids (ions occupy fixed positions and can't move)	Does not conduct electricity in an aqueous solution (unless reacted with water and forms ions)
In molten state (solid converted to liquid by heat) or aqueous solution, they do conduct electricity	In solid, they are soft

Properties of Ionic Substances

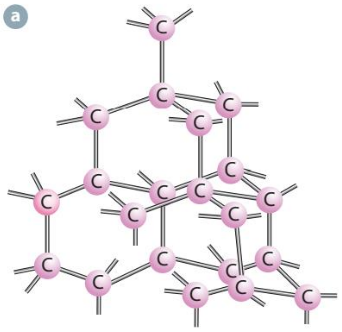
- Lattice structures consisting of opposite charged ions held by electrostatic attraction
- Ionic compounds are represented by empirical formulas (simplified ratio of atoms present)



Properties of Covalent Network Solids

- Consists entirely of covalently bonding atoms.
- High in melting and boiling points
- Insoluble in water
- Hard and brittle
- Non-conductors of electricity in solid/liquid state
- Atoms joined by strong covalent bonds

Covalent Network Structure:

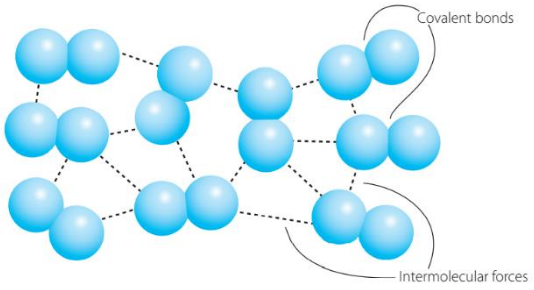


Properties of Covalent Molecular Substances

- Bonding forces holding atoms within the molecule are strong, but forces between one molecule and another molecule are weak (intermolecular forces)

Example:

Boiling separates molecules from one another. Boiling overcomes intermolecular forces and doesn't break any covalent bonds. Thus, because the intermolecular forces are weak, it has a low boiling point.



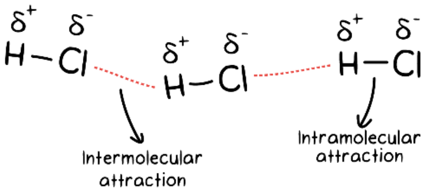
Properties of Metallic Structures

- Density is high
- Good conductors of heat/electricity
- Malleable (can be beaten into sheets)
- Ductile (drawn into wire)
- Shiny surface
- High boiling points

- Consists of a three-dimensional array of positive ions held together by a 'sea' of delocalised electrons
- Due to free movement of electrons, they are shared by positive ions and thus provides the chemical bonding keeping them together

Intermolecular and Intramolecular Bonds

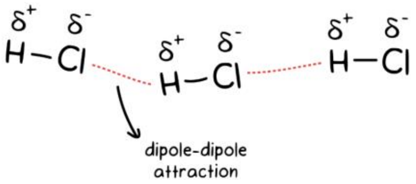
- *Intermolecular forces* are forces existing between molecules
- *Intramolecular forces* are forces which hold atoms together within a molecule
- Intramolecular forces are much stronger than intermolecular forces



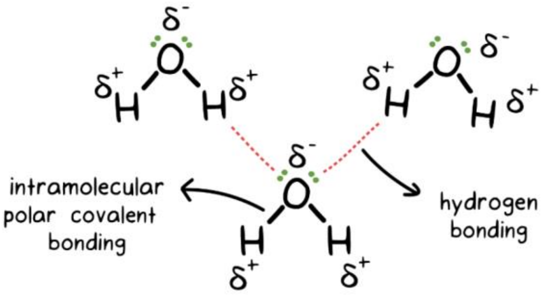
Intermolecular Bonds

Determines physical properties of molecules e.g. boiling point, melting point

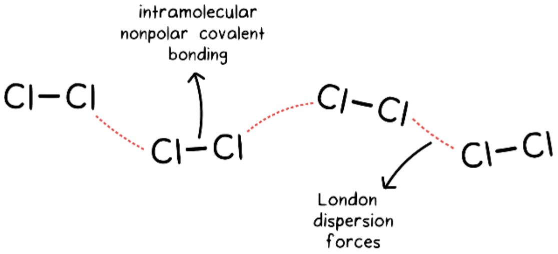
- Dipole-dipole interactions: Occurs when a partially positive part of a molecule reacts with a partially negatively part of a neighbouring molecule.

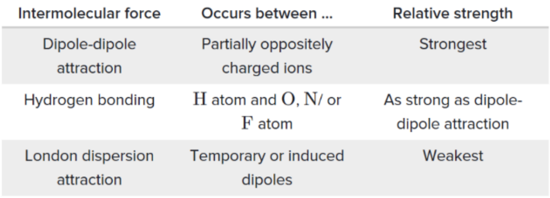


- Hydrogen bonding: Dipole-dipole interaction between a hydrogen atom bonded to an oxygen, nitrogen, or fluorine atom. Strong force of attraction, explains high boiling points/melting points of compounds like water, H_2O, HF



- Dispersion forces: Exists between all kinds of molecules, the more electrons the stronger the force is. E.g. Br_2 has more electrons than Cl_2 so Br_2 has stronger dispersion forces and thus has a higher boiling point





Intermolecular force	Occurs between ...	Relative strength
Dipole-dipole attraction	Partially oppositely charged ions	Strongest
Hydrogen bonding	H atom and O, N/ or F atom	As strong as dipole-dipole attraction
London dispersion attraction	Temporary or induced dipoles	Weakest

The Effect of Bonds on Physical Properties

Boiling and Melting Points:

- Metalloids exhibit covalent network bonding, allowing for high boiling/melting point
- Non-metals have low boiling/melting point (forms covalent substances) because molecules are held together by weak dispersion forces
- Ionic substances have strong electrostatic bonds, thus having high melting and boiling point

Solubility:

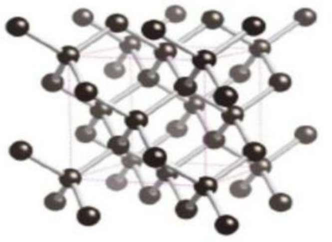
- Polar substances dissolve in polar fluids and vice versa for non-polar
- Strong metallic bonds in metals makes metals almost unable to dissolve in both polar/non-polar fluid
- Strong covalent networks are highly resistant to being dissolved
- Covalent molecular elements are slightly susceptible to being dissolved – each molecule weakly held to others by intermolecular forces
- Ionic substances are extremely soluble in water (polar dissolving in polar)

Allotropy

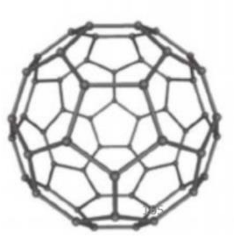
- Different structural forms of an element in the same physical state

Allotropes of Carbon:

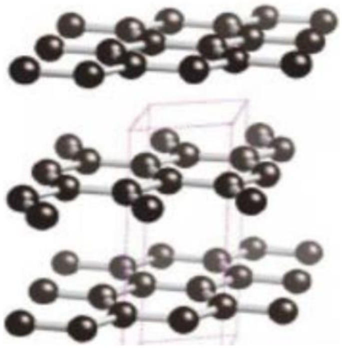
Diamond - Consists of C atoms bonded to 4 other C atoms, does not conduct electricity since all valence electrons are used up in strong covalent bonds



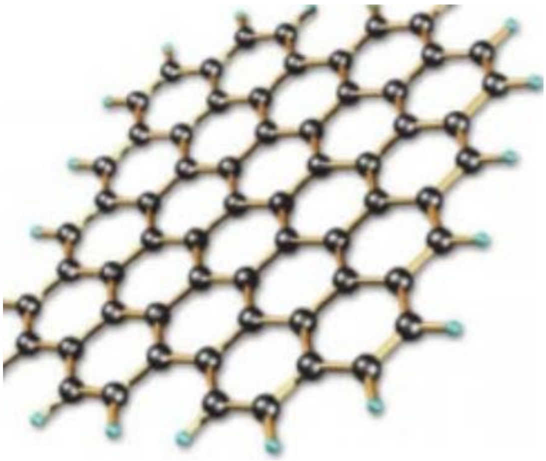
Buckminsterfullerene – Resembles a soccer ball, conducts electricity since each carbon atom has one electron unbonded (forms sea of delocalised electrons)

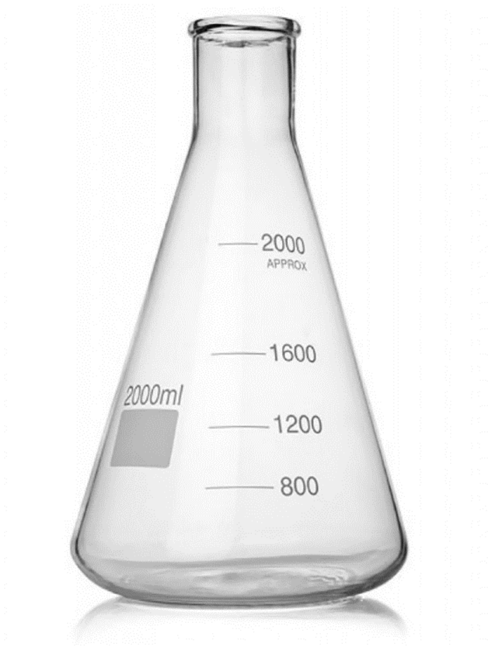


Graphite – Covalent network, each C atom bonded to three other C atoms, leaves one electron unbonded and thus Graphite conducts electricity. Has weak dispersion forces between layers – can slide off easily (for example the use in pencils)



Graphene – Single layer of carbon atoms arranged in a lattice which conducts electricity

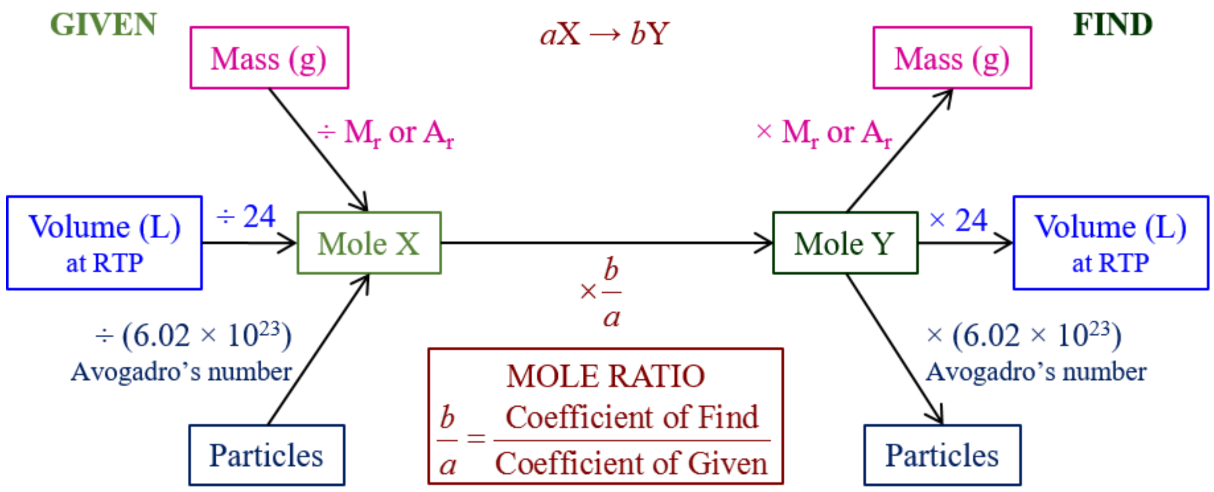




YEAR 11 CHEMISTRY – MODULE 2 NOTES

Introduction to Quantitative Chemistry

STOICHIOMETRY MAP



TOPIC 5: CHEMICAL REACTIONS AND STOICHIOMETRY

Law of Conservation of Mass

- Atoms are neither created nor destroyed
- Mass is conserved; mass of products = mass of reactants

Balancing Chemical Equations

- Ensure number of atoms of each element is the same on both sides (mass/atom conservation)
- Coefficients added to balance equation

STEPS FOR BALANCING CHEMICAL EQ'S:

Example: Hydrogen reacts with oxygen to form water

Step 1 – Word Equation:

Hydrogen + oxygen → water

Step 2 – Chemical Formula Equation:

$H_2 + O_2 \rightarrow H_2O$

Step 3 – Add coefficients

(check atoms of reactants = products):

$H_2 + O_2 \rightarrow 2H_2O$ (balancing oxygen)

$2H_2 + O_2 \rightarrow 2H_2O$ (balances hydrogen)

Step 4 – Add the physical states:

$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{\,(l)}$

Physical States in Chemical Reactions

(*s*) = solid → Usually ionic compounds (salts/solids at room temperature)

(*l*) = liquid (H_2O, ethanol etc.)

(*g*) = gas (all noble gases and some molecules like O_2, CO_2)

(*aq*) = aqueous → a substance is dissolved in water

Relative Atomic Mass, Molecular Mass, Formula Mass

- *Relative Atomic Mass* - Average mass of atoms in naturally occurring element relative to the mass of a Carbon-12 atom. (Refer to Notes on Topic 2: Atomic Structure for example question)

- *Molecular Mass* - Mass of a molecule of a compound relative to the mass of an atom of Carbon-12

Example: Calculate the relative molecular mass of table sugar, sucrose, which has the formula $C_{12}H_{22}O_{11}$.

$$M_M\,(C_{12}H_{22}O_{11}) = (12 \times A_C) + (22 \times A_H)$$

$$+ (11 \times A_O)$$

$$= (12 \times 12.011) + (22 \times 1.008) + (11 \times 15.999)$$

$$= 342.2 \text{ amu}$$

Relative Formula Mass

- Ionic compounds exist as lattice structures, not discrete molecules
- Formula describes ratio → ions present in compound
- RFM → sum of relative atomic masses of atomic species

Example: Calculate the relative formula mass of calcium phosphate, $Ca_3(PO_4)_2$, commonly called rock phosphate and used to make superphosphate fertiliser.

$$F_M\,(Ca_3(PO_4)_2) = (3 \times A_{Ca}) + [2 \times (A_P + (4 \times A_O))]$$

$$= (3 \times 40.08) + [2 \times (30.97 + (4 \times 16.00))]$$

$$= 310.2 \text{ amu}$$

TOPIC 6: MOLE CONCEPT

Avogadro's Number and the Mole

- Atoms or molecules can be grouped into a "mole"
- 1 mole of any substance contains 6.022×10^{23} atoms
- Mole quantities should be stated as atoms, molecules, ions, or formula units.

Number of Particles and Moles

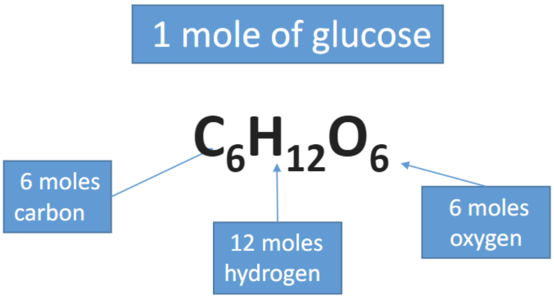
$$\text{number of moles} = \frac{\text{number of particles}}{\text{number of particles in 1 mole}}$$

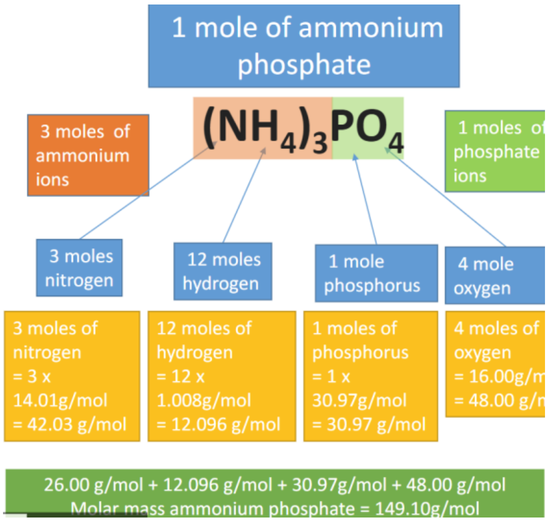
Molar Mass

Units are gmol^{-1} or g/mol

E.g. 1 mole of iron atoms = 55.845g of iron (atomic weight)

- In an equation: subscripts represent how many atoms of each element are present in a molecule of the substance, or the number of moles





Moles and Mass

$$\text{Number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

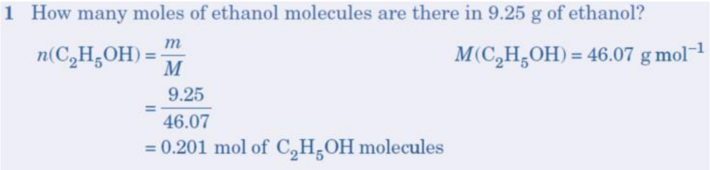
where ***n*** is the number of moles (mol)

$$n = \frac{m}{MM}$$

m is the mass (g)

MM is the molar mass ($\text{g} \cdot \text{mol}^{-1}$)

Example:



1 How many moles of ethanol molecules are there in 9.25 g of ethanol?

$$n(C_2H_5OH) = \frac{m}{M} \qquad M(C_2H_5OH) = 46.07\ \text{g mol}^{-1}$$

$$= \frac{9.25}{46.07}$$

$$= 0.201 \text{ mol of } C_2H_5OH \text{ molecules}$$

Empirical and Molecular Formula

- Empirical Formula – the simplest ratio of atoms present in a compound
- Molecular Formula – the actual ratio of atoms present in a compound

Example:

Empirical formula for Glucose: CH_2O

Molecular formula for Glucose: $C_6H_{12}O_6$

Determining the Empirical and Molecular Formula

- Empirical Formula can be found by using percent composition – dividing the percent of all elements present by its respective molar mass. Then divide by lowest number to simplify ratio and round up.

Example: In a sample of glucose, its composition was indicated to be 40.0% carbon, 6.6% hydrogen, and 53.3% oxygen. Calculate the empirical formula.

n* = *m/MM

$n(C) = 40.0/12.001 = 3.33$

$n(H) = 6.6/1.008 = 6.55$

$n(O) = 53.3/16.00 = 3.33$

Dividing by the simplest ratio:

$n(C) = 3.33/3.33 = 1$

$n(H) = 6.55/3.33 = 2$

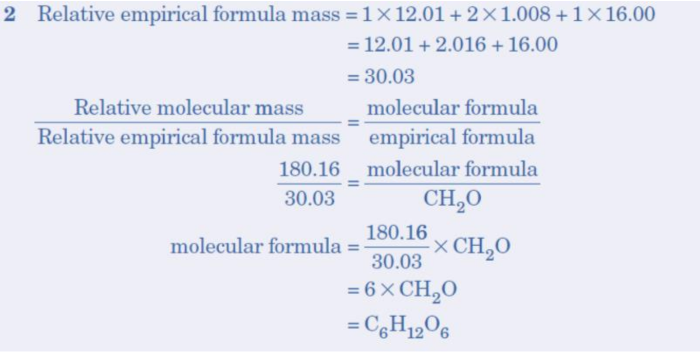
$n(O) = 3.33/3.33 = 1$

The empirical formula is therefore:

CH_2O

[Molecular Formula Example]

Given that the relative molecular mass of glucose is 180.16, what is the molecular formula?



2 Relative empirical formula mass $= 1 \times 12.01 + 2 \times 1.008 + 1 \times 16.00$

$$= 12.01 + 2.016 + 16.00$$

$$= 30.03$$

$$\frac{\text{Relative molecular mass}}{\text{Relative empirical formula mass}} = \frac{\text{molecular formula}}{\text{empirical formula}}$$

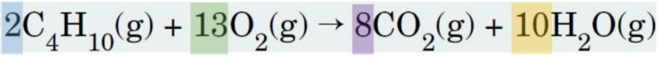
$$\frac{180.16}{30.03} = \frac{\text{molecular formula}}{CH_2O}$$

$$\text{molecular formula} = \frac{180.16}{30.03} \times CH_2O$$

$$= 6 \times CH_2O$$

$$= C_6H_{12}O_6$$

Chemical Equations/Reacting Mass



$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

For example, in this equation, 2 mol of C_4H_{10} react with 13 mol O_2 to form 8 mol CO_2 and 10 mol H_2O

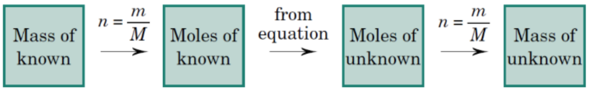
- Using the coefficients, it can be determined how much of a substance is needed or produced.

Example: If 4.5 mol of butane is burnt in oxygen, calculate the number of moles of oxygen required to completely react with the butane.

$$\frac{13 \text{ mol } O_2}{2 \text{ mol } C_4H_{10}} \times 4.5 \text{ mol } C_4H_{10} = \frac{13 \text{ mol } O_2 \times 4.5}{2}$$

$$= 29 \text{ mol } O_2$$

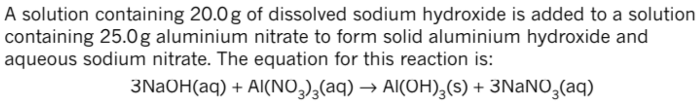
In carrying out mass–mass calculations it is useful to remember the following method.



Limiting Reagents

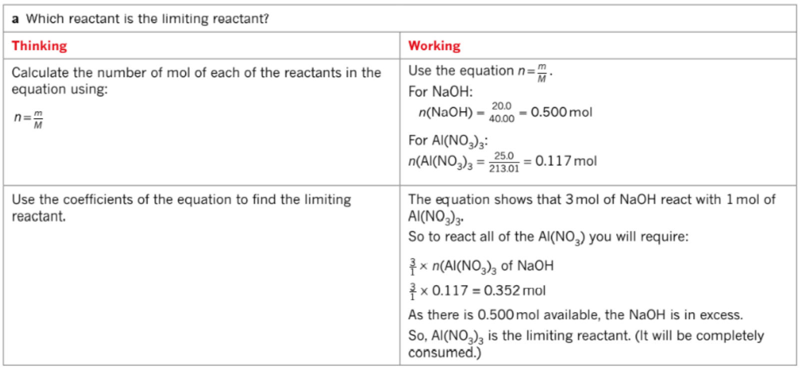
- The *limiting reagent* in a reaction is the reactant which is used up first/completely consumed.
- The *excess reactant* is the reactant which is not completely consumed.

[Example of Limiting Reagent Question]



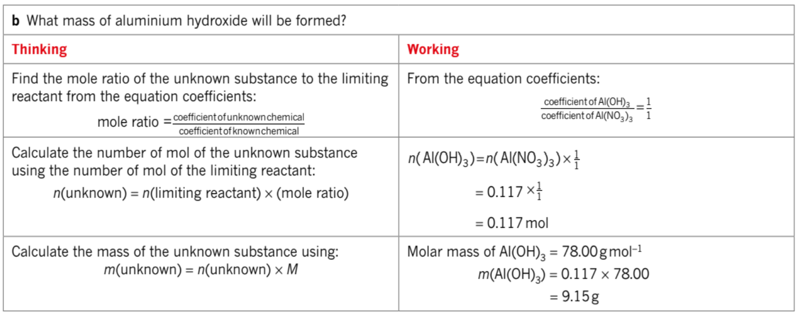
A solution containing 20.0 g of dissolved sodium hydroxide is added to a solution containing 25.0 g aluminium nitrate to form solid aluminium hydroxide and aqueous sodium nitrate. The equation for this reaction is:

$$3NaOH(aq) + Al(NO_3)_3(aq) \rightarrow Al(OH)_3(s) + 3NaNO_3(aq)$$



a Which reactant is the limiting reactant?

Thinking	Working
Calculate the number of mol of each of the reactants in the equation using: $n = \frac{m}{M}$	Use the equation $n = \frac{m}{M}$. For NaOH: $n(NaOH) = \frac{20.0}{40.00} = 0.500$ mol For $Al(NO_3)_3$: $n(Al(NO_3)_3 = \frac{25.0}{213.01} = 0.117$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows that 3 mol of NaOH react with 1 mol of $Al(NO_3)_3$. So to react all of the $Al(NO_3)$ you will require: $\frac{3}{1} \times n(Al(NO_3)_3$ of NaOH $\frac{3}{1} \times 0.117 = 0.352$ mol As there is 0.500 mol available, the NaOH is in excess. So, $Al(NO_3)_3$ is the limiting reactant. (It will be completely consumed.)



b What mass of aluminium hydroxide will be formed?

Thinking	Working
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: mole ratio $= \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$	From the equation coefficients: $\frac{\text{coefficient of } Al(OH)_3}{\text{coefficient of } Al(NO_3)_3} = \frac{1}{1}$
Calculate the number of mol of the unknown substance using the number of mol of the limiting reactant: $n(\text{unknown}) = n(\text{limiting reactant}) \times (\text{mole ratio})$	$n(Al(OH)_3) = n(Al(NO_3)_3) \times \frac{1}{1}$ $= 0.117 \times \frac{1}{1}$ $= 0.117$ mol
Calculate the mass of the unknown substance using: $m(\text{unknown}) = n(\text{unknown}) \times M$	Molar mass of $Al(OH)_3 = 78.00 \text{ g mol}^{-1}$ $m(Al(OH)_3) = 0.117 \times 78.00$ $= 9.15$ g

TOPIC 7: CONCENTRATION AND MOLARITY

Units of Concentration

- g/L or gL^{-1}
  (mass of solute per litre of solution)
- mol/L or $molL^{-1}$
  (moles of solute per litre of solution)
- parts per million (ppm – mg/kg)
- percentage by mass

[Example in gL^{-1}]

What is the concentration in gL^{-1} of a solution containing 5.00g of glucose in 250mL of solution?

250mL = 0.25L

Therefore:

Concentration = 5.00g/0.25L = $20.0gL^{-1}$

[Example in ppm (parts per mil)]

A sample of tap water was found to contain 0.0537g of sodium chloride (NaCl) per 250.0g of solution. Calculate the concentration of NaCl in parts per million (ppm).

0.0537g = 53.7mg, 250.0g = 0.25kg

Therefore:

Concentration = 53.7mg/0.25kg = 215ppm

Percentage by mass - %(w/w):

Describes mass of solute present in 100g of solution. For example, a concentration of 0.9%(w/w) has 0.9g of solute per 100g of solution.

Percentage by volume - %(v/v):

Describes volume of solute present in 100mL of solution.

$$\text{concentration } (\%\text{v/v}) = \frac{\text{volume of solute (in mL)}}{\text{volume of solution (in mL)}} \times 100\%$$

Calculating Molarity

$$c = \frac{n}{v}$$

c = concentration (in mol L^{-1})

n = amount in moles (mol)

v = volume in litres

[Example]

Calculate the molar concentration of a solution that contains 0.24mol of glucose dissolved in 500mL of solution.

500mL = 0.5L

Therefore:

Molarity = 0.24mol/0.5L = 0.48mol L^{-1}

Dilutions

- Diluted solutions contain a lot of water (such as adding water to cordial)
- Dilution doesn't change amount of solute, only the concentration decreases

$$c_1v_1 = c_2v_2$$

c = concentration (mol L^{-1})

v = volume

[Example]

Calculate the concentration of the solution formed when 10.0mL of water is added to 5.00mL of 1.2mol L^{-1} HCl.

c_1 = 1.2 mol L^{-1}

v_1 = 5.00 mL

v_2 = 10.0mL + 5.0mL = 15.0mL

$c_2 = c_1\, v_1 / v_2$

= 1.2mol L^{-1} * 5.00mL / 15.0 mL

= 0.40 mol L^{-1}

TOPIC 8: GAS LAWS

Nature of Gases

- States of matter with no fixed shape/volume
- Lower density than solids or liquids
- Great deal of empty space between particles
- Particles have a lot of kinetic energy
- Particles diffuse and spread out evenly throughout the whole volume of area

Kinetic Molecular Theory

- Gases are composed of small particles (atoms or molecules)
- Volume of particles is negligible compared with volume they occupy
- Gas particles move in random straight-line motion
- Forces between particles are negligible
- Average kinetic energy of a particle increases as gas is increased

Nature of Volume

- Gases have volume of a container because they occupy all space

Units:

1mL = $1cm^3$

1L = 1000mL

$1m^3 = 1 \times 10^6 cm^3$

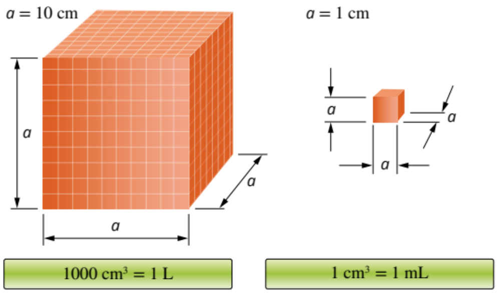


FIGURE 9.1.5 Cubic centimetres, millilitres and litres are the most commonly used units of volume.

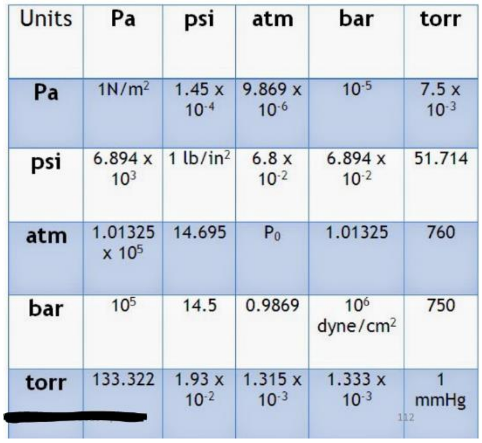
Nature of Pressure

Pressure is defined as the amount of force exerted per area:

$$P = \frac{F}{A}$$

- Rapid motion and collisions of molecules with walls of a container causes pressure.

Units of Pressure:

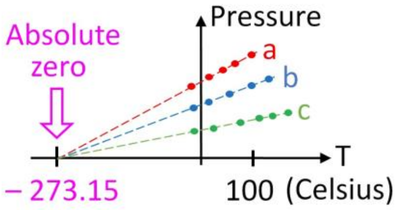


Units	Pa	psi	atm	bar	torr
Pa	$1N/m^2$	1.45×10^{-4}	9.869×10^{-6}	10^{-5}	7.5×10^{-3}
psi	6.894×10^3	$1 lb/in^2$	6.8×10^{-2}	6.894×10^{-2}	51.714
atm	1.01325×10^5	14.695	P_0	1.01325	760
bar	10^5	14.5	0.9869	10^6 dyne/cm^2	750
torr	133.322	1.93×10^{-2}	1.315×10^{-3}	1.333×10^{-3}	1 mmHg

[redacted]

Nature of Temperature

- Temperature – average measure of kinetic energy in a system
- Commonly used scales include Celsius (°C), Fahrenheit (°F) and Kelvin (K).
- Coldest theoretical temperature is absolute zero – thermal motion of particles in matter reaches a minimum.



Boyle's Law

$$PV = k$$

Where P is the pressure, V is the volume, and k is a constant.

- If original volume and pressure is known, the volume of a gas can be worked out with a new given pressure:

$$P_1V_1 = P_2V_2$$

- Increase in pressure = decrease in volume, and vice versa.

Charles' Law

- Increase in temperature of volume of gases causes individual gas molecules to move faster

$$\frac{V}{T} = k$$

Where V is volume, T is temperature in kelvin (K)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

CONVERTING FROM DEGREES TO KELVIN:

- Add degrees given to 273.15°

Example: 23° = 23 + 273.15 = 296.15K

Gay-Lussac's Law

- Pressure of a given amount of gas held at constant volume is directly proportional to the Kelvin temperature

$$\frac{P}{T} = k$$

Where P is pressure, T is temperature in kelvin (K)

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Avogadro's Law

$$\frac{V}{n} = k$$

Where V is the volume, n is the moles, k is a constant.

Thus:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Combined Gas Law

- Shows relation between pressure, volume, and temperature (k is a constant)

$$\frac{PV}{T} = k$$

Thus:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Ideal Gas Law

- This law takes all the gas laws into account.

$$PV = nRT$$

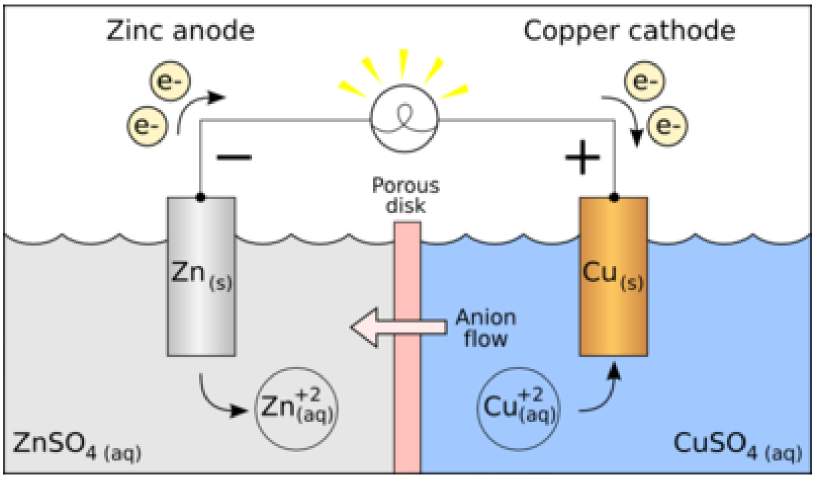
P is the pressure

V is the volume

n is the number of moles

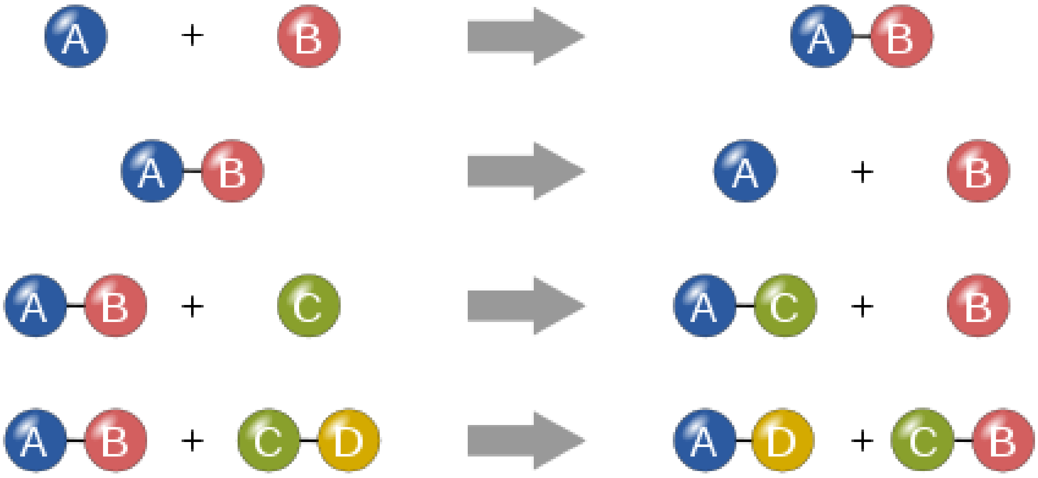
R = 8.314 J mol^{-1} K^{-1} or 8.314 L kPa mol^{-1} K^{-1}

T is the temperature in Kelvin (K)



YEAR 11 CHEMISTRY – MODULE 3 NOTES

Reactive Chemistry



TOPIC 9: CHEMICAL REACTIONS

Indicators of Chemical Changes

- Bubbles
- Colour Change
- Odour Change
- Precipitate Formed
- Release or absorption of energy

Types of Chemical Reactions

Synthesis	Two or more reactions combine to form a single product. $A + B \rightarrow AB$ Example: $2H_2 + O_2 \rightarrow 2H_2O$
Decomposition	One reactant breaks down into two products. $AB \rightarrow A + B$ Example: $CaCO_3 \rightarrow CO_2 + CaO$
Combustion	Involving oxygen, reactions where a reactant is burning. Molecules containing carbon, hydrogen, oxygen will burn in oxygen to form water and another carbon-based product. Hydrocarbon + $O_2 \rightarrow H_2O + CO_2$ $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
Precipitation (Double Displacement)	Cations and anions in aqueous solutions combine to form a precipitate (insoluble ionic solid). This reaction can be determined by the solubility rules. $AB + CD \rightarrow AD + CB$ $CdSO_4 + K_2S \rightarrow CdS + K_2SO_4$

Solubility Rules

- Whether or not a precipitate is formed is dictated by solubility rules.
- Provides guidelines that tell which ions forms solids and which remain in their ionic form in aqueous solution.
- Rules are to be followed from top down, if something is insoluble or soluble due to rule 1, has precedence over a higher numbered rule.

SOLUBLE	***INSOLUBLE***
- Group 1 cations and NH_4^+	- Salts containing silver, lead, mercury (I)
- Acetates, nitrates, and perchlorates	- Carbonates, phosphates, sulphides, oxides, hydroxides
- Bromides, iodides, and chlorides	- Sulfides formed with group 2 cations and hydroxides formed with calcium, strontium, and barium are exceptions.
- Sulfates with exception of sulfates formed with calcium, strontium, barium	

If an ion is soluble → remains in aqueous ion form

If ion is insoluble → forms a solid with an ion from the other reactant

If all ions in reaction are soluble → no precipitation reaction occurs.

Net Ionic Equations

In a precipitation reaction:

$$AB_{(aq)} + CD_{(aq)} \rightarrow AD_{(aq)} + CB_{(s)}$$

To write a net ionic equation, first separate soluble products into their ions:

$$A^+ + B^-_{(aq)} + C^+ + D^-_{(aq)} \rightarrow A^+ + D^-_{(aq)} + CB_{(s)}$$

Since A^+ and D^- are on both sides they can be eliminated (spectator ions – watching the reaction happen).

Thus:

$$C^+_{(aq)} + B^-_{(aq)} \rightarrow CB_{(s)}$$

Acid-base Reactions

pH Theory:

pH of 0-7 → acidic

pH of 7→ neutral

pH of 7-14 → basic/alkaline

- Acids → substances which in a solution produce H^+ ions.
- In aqueous solutions, *hydrogen ions exist as hydronium ions* (H_3O^+)
- Bases → Contain OH^- or O^{2-}

NEUTRALISATION REACTIONS:

General Form:

acid + base → salt(ionic) + water

Properties of Acids and Bases

- *Acids are proton* (H^+) *donors* – forms hydronium ions in water. For example:

$$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

- *Bases are proton* (H^+) *acceptors* – forms hydroxide ions (OH^-) in water.

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

Gain of proton (H^+)

Loss of proton (H^+) forming hydroxide ion

BASIC AND ACIDIC OXIDES:

- Oxides which *dissolve* and *react* with ***strong acids*** are basic oxides.
- Oxides which *dissolve* and *react* with ***strong basic solutions*** are acidic oxides.

Acid-Carbonate Reactions

When acids reaction with metal carbonates:

Acid + metal carbonate → salt + H_2O + CO_2

Example of acid/carbonate reaction:

Magnesium carbonate + hydrochloric acid → magnesium chloride + water + carbon dioxide

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$$

Chemical Processes of Detoxifying Fruits by Aboriginal and Torres-strait Islanders

OBSERVATIONS OF LEACHING OF TOXINS IN BUSH TUCKER:

Process of detoxification:

- Bake seeds for 6 hours with hot coal, cover in leaves and earth.
- Cut seeds into small pieces
- Place in woven bag and submerge in a creek

Chemical processes:

→ Breaks down starch

→ Increases surface area in contact with water

→ Allows water to percolate through

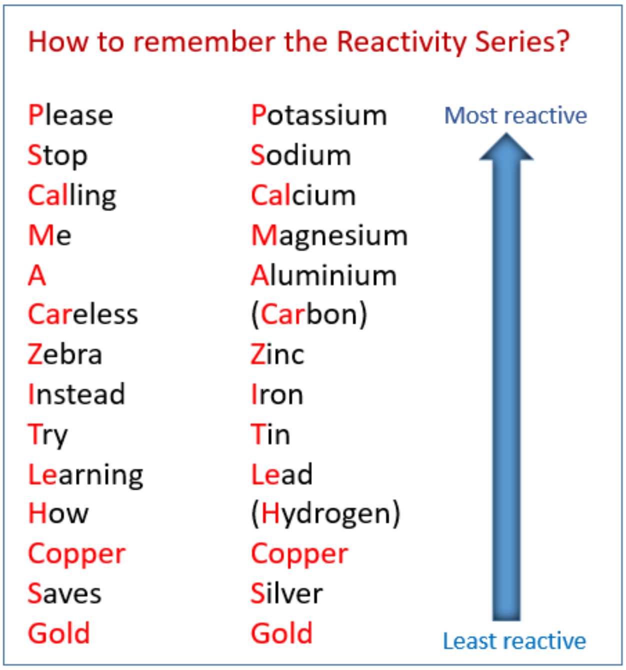
chopped seeds and removes toxic substances.

TOPIC 10 – PREDICTING REACTIONS OF METALS

Reactions of Metals

Substance	***Reaction***
Water	• Most metals don't react with cold water If a metal reacts with water, it forms hydrogen, and a metal hydroxide. e.g. sodium + water → hydrogen + sodium hydroxide $2Na + 2H_2O \rightarrow 2NaOH + H_2$
Acids	• Metals that don't react with water *may* react with dilute acids. If a metal reacts with an acid, it forms a salt + hydrogen e.g. zinc + hydrochloric acid → zinc chloride + hydrogen $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
Oxygen	Most metals react with oxygen in air to form metal oxides. e.g. iron + oxygen → iron (III) oxide $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
Metals with other metallic ions in solution (More reactive metal displaces the lesser reactive metal from solution)	Known as metal displacement reactions. Example: Solid zinc is placed into solution of copper sulfate (blue in colour). When this happens, two electrons are taken from the zinc by the copper (copper gains two electrons). Thus, the copper becomes neutral and forms solid copper in the solution. Solution becomes zinc sulfate which is clear in colour. Equation: $Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4\,(aq)$ $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$

Metal Activity Series



How to remember the Reactivity Series?

Please	Potassium	Most reactive
Stop	Sodium	
Calling	Calcium	
Me	Magnesium	
A	Aluminium	
Careless	(Carbon)	
Zebra	Zinc	
Instead	Iron	
Try	Tin	
Learning	Lead	
How	(Hydrogen)	
Copper	Copper	
Saves	Silver	
Gold	Gold	Least reactive

The metal reactivity series lists metals in order of highly reactive to very low reactivity.

FIRST IONISATION ENERGY:

- Highly reactive metals have low first-ionisation energy
- These metals are highly reactive – they want to lose an electron.

Example: $K(s) \rightarrow K^+ + e^-$, $Ca(s) \rightarrow Ca^{2+} + 2e^-$

ATOMIC RADII:

- Highly reactive metals have large atomic radii.
- Outer electrons not held tightly by the nucleus and are therefore free to react with other elements

ELECTRONEGATIVITY:

- Highly reactive metals have low electronegativity, they are happy to give away electrons

Redox Reactions ~ Half Equations

- Redox reactions (reduction-oxidation reactions) – reactions involving transfer of electrons from one substance to another

Mnemonic to help remember the process:

OIL RIG

Oxidation Is Loss →
Reduction Is Gain →
of electrons (e^-)

Examples:

The reaction of Magnesium metal and Oxygen:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Magnesium loses two electrons; hence it becomes Mg^{2+} in ionic MgO. Thus, Mg is being *oxidised*. Oxygen has lost two electrons to become an O^{2-} ion, therefore Oxygen has been reduced.

HALF EQ'S WITH REDOX:

Half equations describe oxidation and reduction processes separately in terms of electrons gained or lost.

Example:

Determine the half equation for:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Zinc loses electrons, oxidation half equation can be represented as:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

Hydrogen gains electrons, reduction half equation can be represented as:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

Redox Reactions ~ Oxidation Numbers

Can be difficult to tell if a reaction is a redox reaction, so there are rules to help to determine oxidation number:

Rule	*Examples*
Elements ***by themselves*** in their elemental state have an oxidation number of 0.	Cl = 0 H_2 = 0 F = 0 O_2 = 0
Monoatomic ions – oxidation number is charge on the ion	Cu^{2+} has an oxidation number of +2. Cl^- has an oxidation number of -1.
Sum of oxidation states of all atoms in neutral molecules or ionic compounds must equal to zero	The sum of oxidation states of H_2O must equal 0. Similarly, for HF etc.
Group 1A is always +1, Group 2A is always +2 ***in compounds***	In the above example, H_2O, H has an oxidation number of +1.
Sum of oxidation states of all atoms in a polyatomic ion equals the charge on the ion.	The sum of oxidation states of all atoms SO_4^{2-} must equal -2.
For combined oxygen, state is usually -2. Exceptions are -1 in peroxides and +2 in F_2O	In the above example for SO_4^{2-}, O must have an oxidation state of -2.

Examples:

What is the oxidation number (state) of sulfur in sulfuric acid, H_2SO_4?

- H is +1 (Group 1A is always +1 in compounds).
- Oxygen is -2 (Combined oxygen is usually -2).
- Molecule is neutral so the sum of oxidation numbers must equal 0:

Let the unknown oxidation number S be *x*.

$$+2 + x + (-8) = 0$$

2 hydrogen atoms multiplied by +1 oxidation number to get +2. 4 oxygen atoms multiplied by an oxidation number of -2 to get -8.

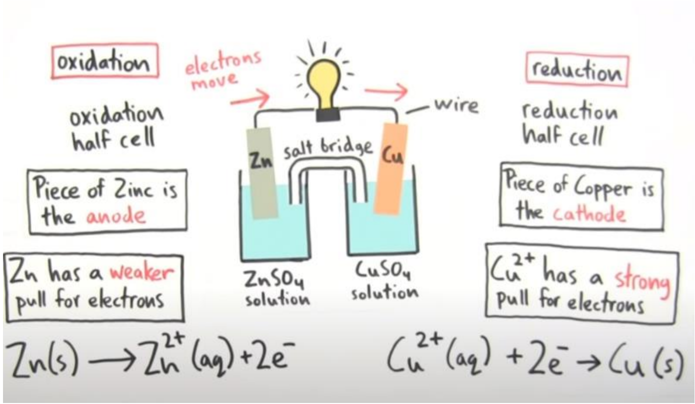
Solving for *x*:

x = +6, thus since *x* = S, the sulfur atom has a +6 ~~charge~~ oxidation number

Redox Reactions ~ Galvanic and Electrolytic Cells

GALVANIC CELLS:

Devices which use chemical reactions to create electricity



COMPONENTS:

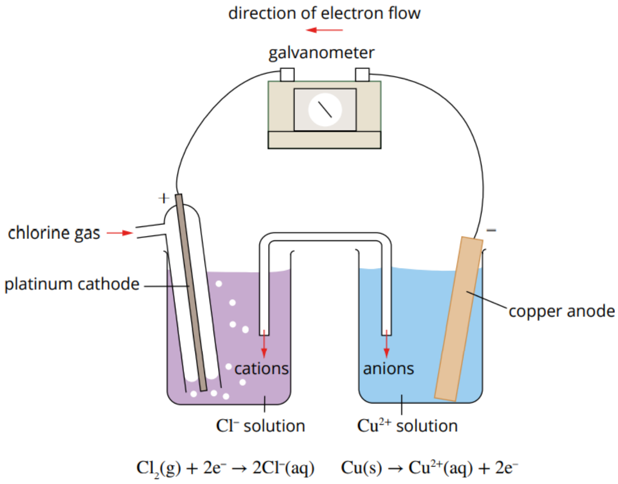
- Anode: Electrode of oxidation
- Cathode: Electrode of reduction
- Salt Bridge: Helps balance out charge on both half cells

HOW DO GALVANIC CELLS WORK?

- For diagram above, Zn was placed in a $ZnSO_4$ solution, and Cu was placed in a $CuSO_4$ solution.
- The Cu^{2+} ions in the copper solution have a *strong* pull for electrons. Hence, for one Cu^{2+} ion, 2 electrons from the zinc metal are transferred over to the reduction half cell (where the copper solution is). This movement of electrons is what causes electricity or an *electrical current.*
- These electrons join up with the Cu^{2+} ion to make it neutral solid Cu, and it then joins up with the rest of the copper metal. In turn, because 2 electrons were lost, a Zn^{2+} ion is ejected from the Zn metal, and dissolves into the solution.

PURPOSE OF THE SALT BRIDGE:

- Contains ions free to move through the cotton wool so they can balance the charges formed in the half cells.
- Without a salt bridge, one half cell would accumulate negative charge, and the other one a positive charge. This will cause the cell to stop working.



STANDARD ELECTRODE POTENTIALS:

- The potential difference of electrodes of the cell when a current is drawn (max. voltage a cell can deliver).

HOW TO CALCULATE:

When reading standard potential chart:

- $E_{\text{total}} = E_{\text{oxidation}} + E_{\text{reduction}}$

Going down the left-hand side of the electrode potential chart – higher chance of being reduced.

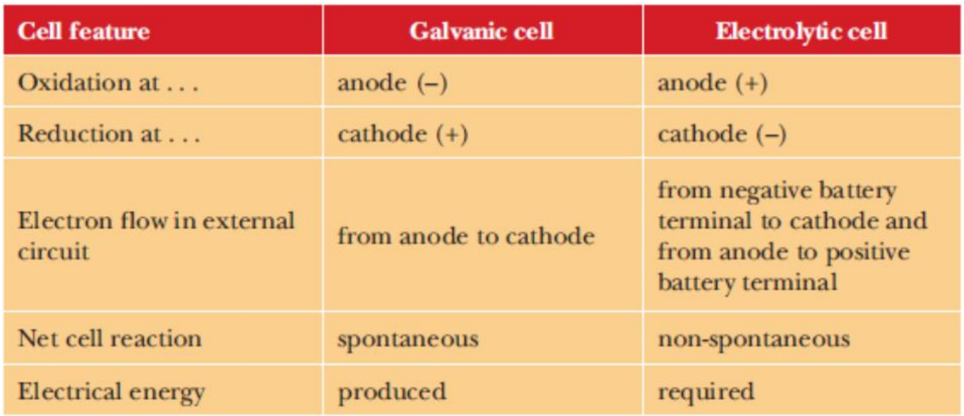
Going up the right-hand side of the electrode potential chart – higher chance of being oxidised.

- Split full equation in two half equations, find the first half equation involved and then go down the list to find the other one (reduced), take note of voltage of that half equation.
- Likewise going up for oxidation, and take note of voltage but swap sign (half equation read other way around – the chart is for reduction not oxidation).
- Add both.

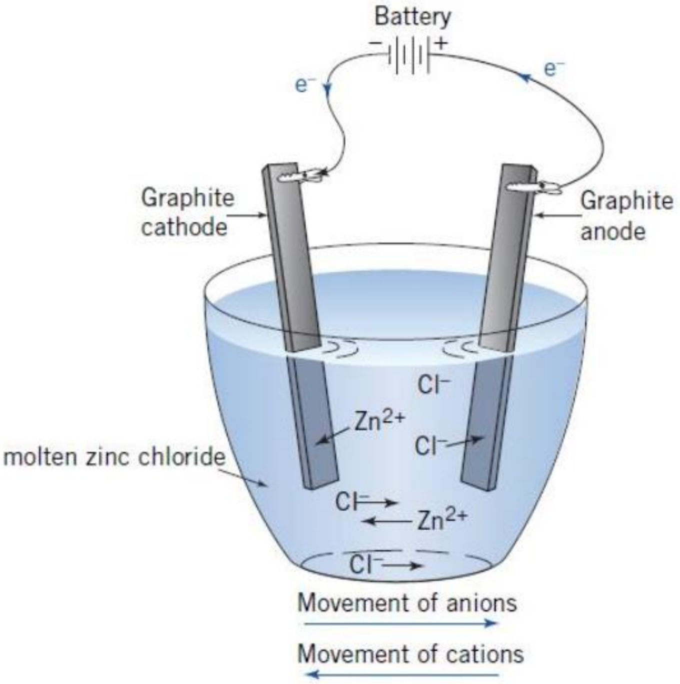
ELECTROLYTIC CELLS

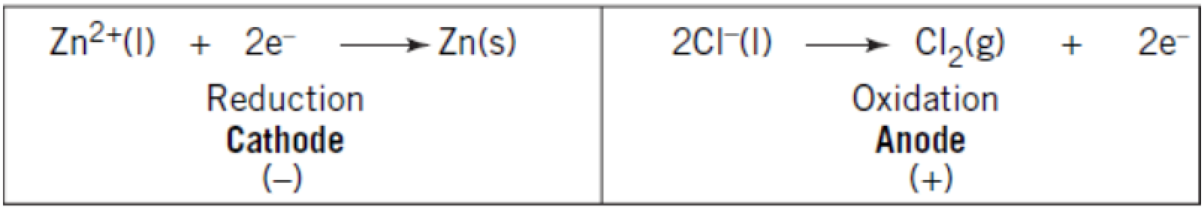
- Opposite of galvanic cells – electricity causes a chemical change (electrical energy is converted to chemical).

ELECTROLYTIC CELLS



Cell feature	Galvanic cell	Electrolytic cell
Oxidation at . . .	anode (–)	anode (+)
Reduction at . . .	cathode (+)	cathode (–)
Electron flow in external circuit	from anode to cathode	from negative battery terminal to cathode and from anode to positive battery terminal
Net cell reaction	spontaneous	non-spontaneous
Electrical energy	produced	required





$Zn^{2+}(l) + 2e^- \longrightarrow Zn(s)$ Reduction **Cathode** (–)	$2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$ Oxidation **Anode** (+)

TOPIC 11 – RATES OF REACTION

Rates of Reaction

- Reaction rate is a measure of the reactants being used or the products being formed per unit time.

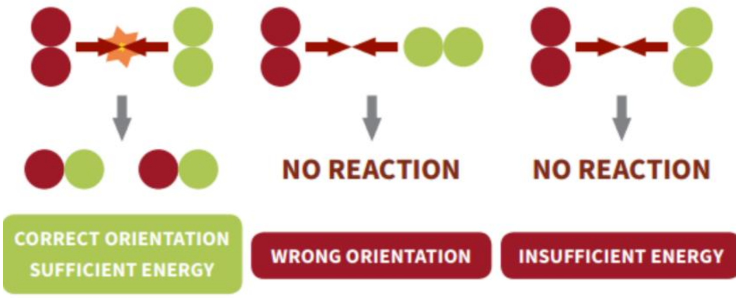
Reaction Rate =

$$\frac{volume\ consumed/produced}{time\ taken}$$

Collision Theory of Reaction Rates

Factors to be considered for collision of molecules:

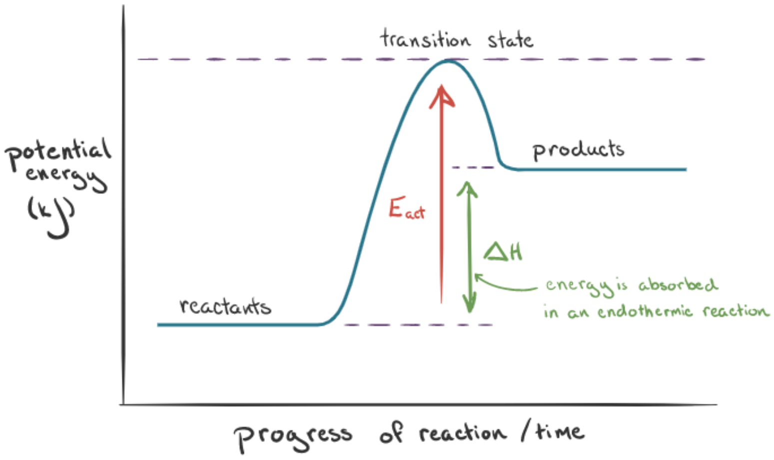
- Two molecules *may* only react when they come into contact.
- Correct Orientation
- Sufficient Energy (activation energy)



ACTIVATION ENERGY:

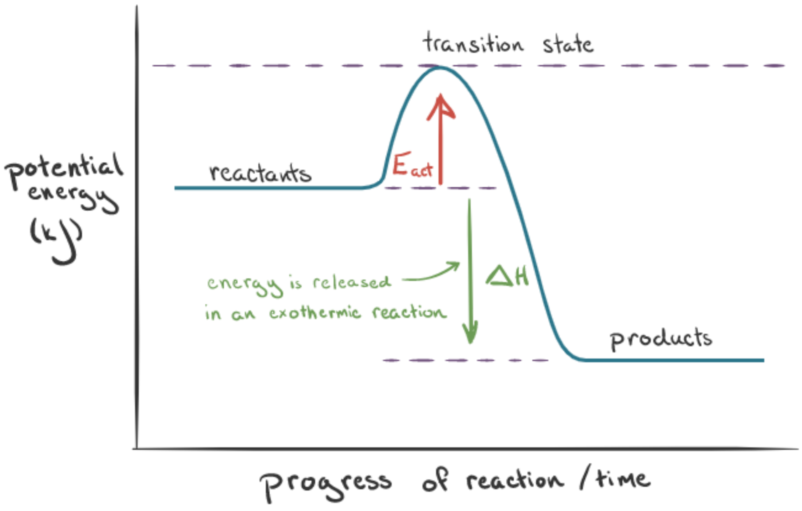
- Species won't react unless they collide with the minimum energy needed (activation energy). This can be shown with an energy profile

ENDOTHERMIC REACTION ENERGY PROFILE



(Heat is absorbed in reaction)

EXOTHERMIC REACTION ENERGY PROFILE



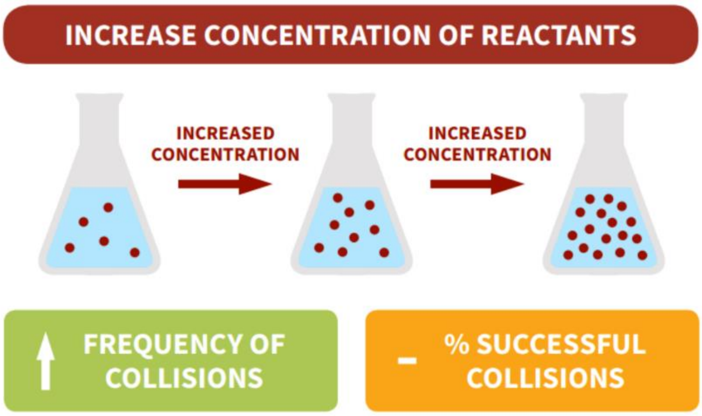
(Heat is released in the reaction)

- If particle collides with less energy than activation energy, nothing happens.

OTHER FACTORS FOR COLLISIONS:

Concentration:

- Increasing concentration increases rate of reaction – greater number of particles able to react, thus increasing frequency of collision between two particles.

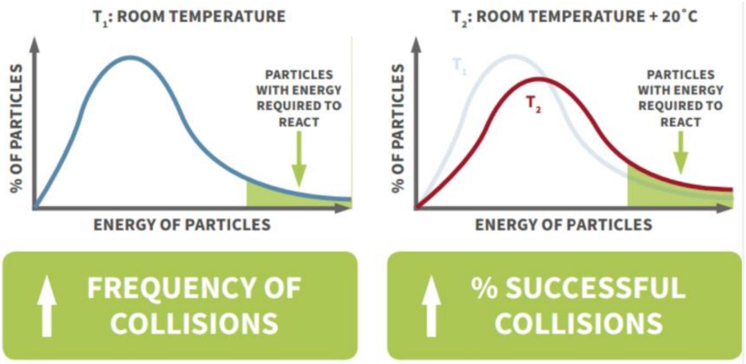


Surface Area:

- Increasing surface area of reactants increases number of particles exposed and are available to react, thus increases frequency of particle collisions, and in turn increases rate.

Temperature:

- Increase in temperature = increase in kinetic energy of particles. Increases frequency of particle collisions and greater proportion of collisions will have energy to react.



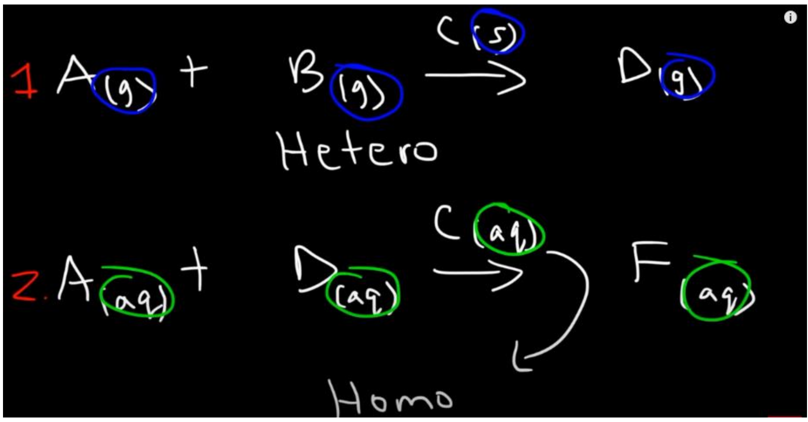
Area under curve represents no. of particles at a given temp.

Catalysts:

- Provide alternative reaction pathways, which reduces activation energy needed to start the reaction. Particle collisions therefore need less energy, increases ROR.
- Catalysts **are not permanently consumed** in a reaction.

HOMOGENOUS CATALYSTS – Catalyst in same phase as reactants/products.

HETEROGENOUS CATALYSTS – Catalyst in different phase to reactants/products. Provides surface to lower activation energy.

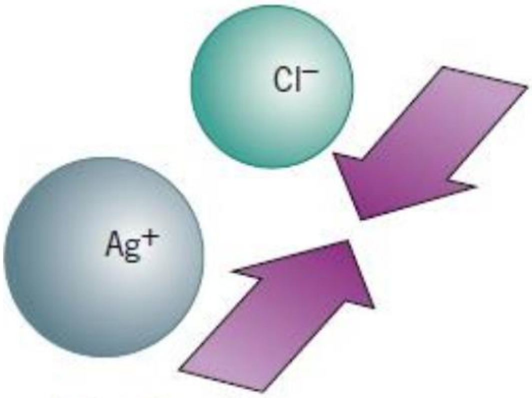


Pressure:

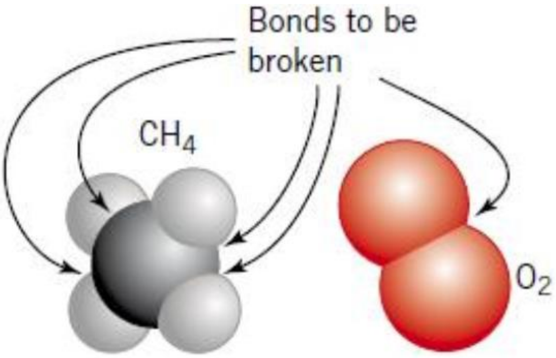
- Increasing pressure of gases forces particles closer together. Increases frequency of particle collisions and therefore increases ROR.

In general – General Guidelines:

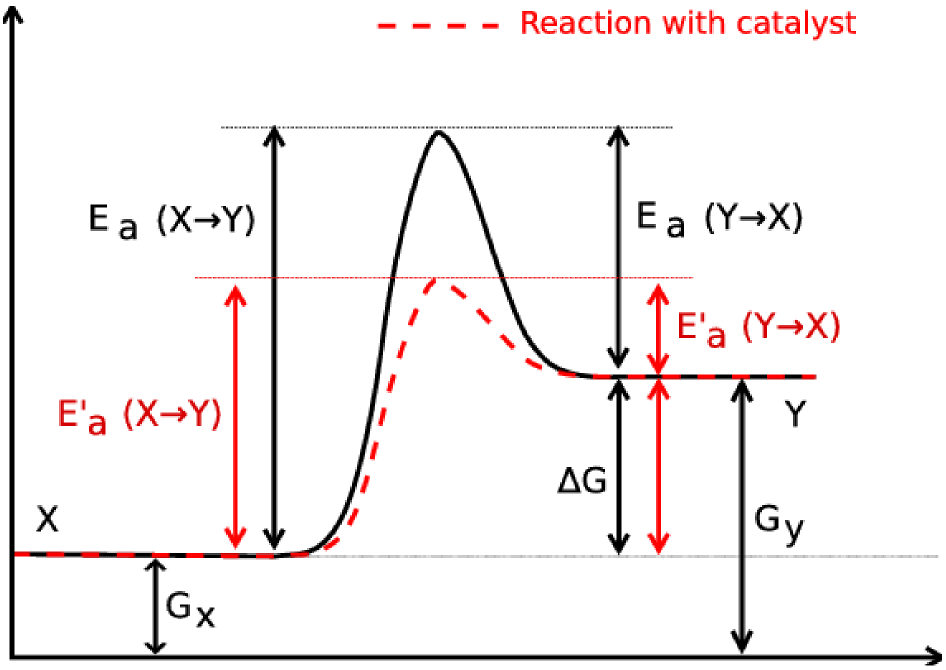
- If a reaction doesn't involve bonding rearrangements, reaction is rapid at room temp.
- If a reaction involves breaking of bonds, slow at room temperature. Process only takes place if there is sufficient energy and orientation of particles.



A simple collision between Ag^+ and Cl^- can cause reaction.



The reaction between CH_4 and O_2 is comparatively slow due to bond-breaking and bond-forming processes.



YEAR 11 CHEMISTRY – MODULE 4 NOTES

Drivers of Reactions

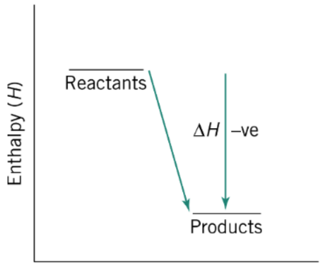
TOPIC 12 – ENERGY CHANGES IN CHEMICAL REACTIONS

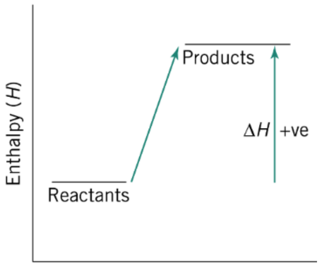
Enthalpy

- Chemical energy stored within a substance is called the *enthalpy* of the substance, enthalpy denoted by ΔH.
- Change of enthalpy calculated rather than enthalpy of a system – can't be measured.
- Atoms or molecules involved in reaction is referred to as a *system* and anything around it is the *surroundings.*

To calculate the heat of reaction (ΔH) from products and reactants:

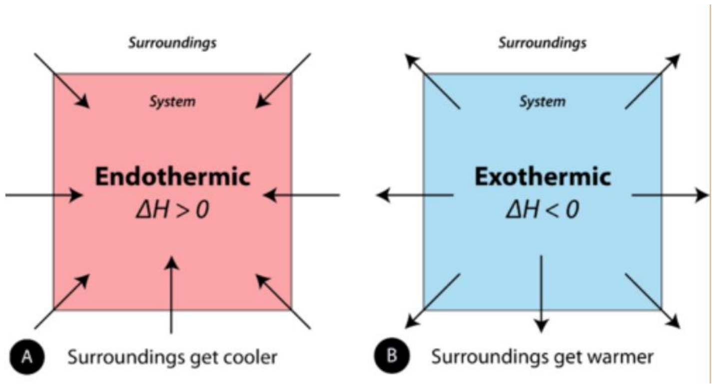
$$\Delta H = \sum H_{products} - \sum H_{reactants}$$



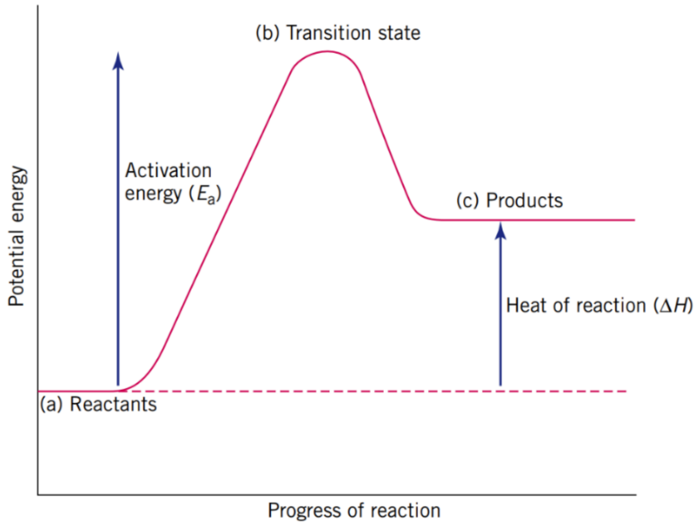


Endothermic and Exothermic Reactions

ENDOTHERMIC	*EXOTHERMIC*
Absorb energy (heat)	Release energy (heat)
Surroundings become cooler. System becomes warmer.	Surroundings become warmer. System becomes cooler.
Happens when ΔH > 0	Happens when ΔH < 0



Energy Profile Diagrams

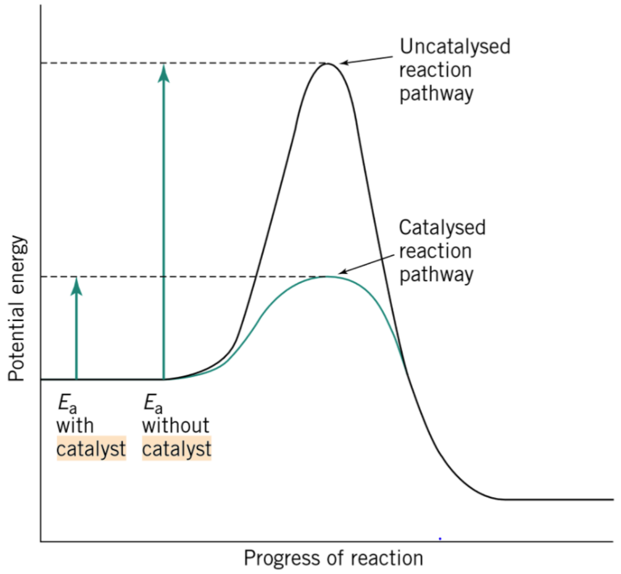


ΔH is difference between reactants and products, if reaction is exothermic then reactants have higher enthalpy, if reaction is endothermic products have higher enthalpy (like the diagram above).

- Transition state – represents highest energy state for reacting system. Bond breaking and formation taking place here.

Catalysts

- Provides an alternate reaction pathway by lowering activation energy. Allows the reaction to occur in such a way that less collision energy is required. Catalysts are not permanently consumed in a reaction.



Calculating Heat Absorbed/Released

Formula for calculating the amount of heat absorbed or released in a reaction is given by:

$$q = mc\Delta T$$

Where:

q = heat energy absorbed/released (in joules, J)

m = mass of substance being heated/cooled (g)

c = specific heat capacity of substance being cooled ($J\ g^{-1}\ K^{-1}$ or $J\ g^{-1}\ °C^{-1}$).

ΔT = temperature change of substance being heated or cooled (°C Celsius or K kelvin).

$$\Delta T = \textit{final temperature} - \textit{initial temperature}$$

Molar Heat of Solution

- Molar heat of solution refers to heat energy absorbed or released when one mole of solute dissolves in solution.
- E.g. in the reaction:

$$NaOH(s) \rightarrow Na^+(aq) + 2OH^-(aq)$$

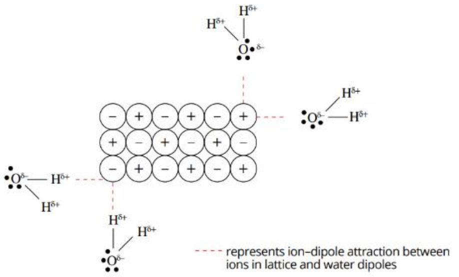
$$\Delta H = -44.5kJ/mol$$

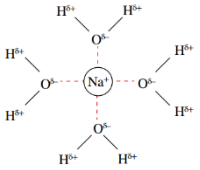
The dissolution of sodium hydroxide *releases* 44.5kJ (negative enthalpy value) of heat energy per mole dissolved.

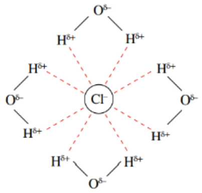
Dissociation of Ionic Substances in Water

1. Electrostatic forces broken between positive and negative ions. Process uses energy – endothermic.
2. Intermolecular forces (hydrogen bonds) are broken, which uses energy – endothermic.
3. Rearranging of bonds between positive and negative ions/water molecules creating an ion-dipole bond. Releases energy and is exothermic.

DIAGRAMS OF DISSOCIATION:







Measuring Molar Heat of Solution

- Measured using a calorimeter – e.g. a polystyrene foam cup. Insulation of these kinds of cups prevents heat transfer to and from surroundings.
- ΔH can be calculated if temperature change is measured and heat capacity of calorimeter and water are known:

$$\Delta H = -\frac{q}{n}$$

Where:

ΔH = Chemical energy released/absorbed (kJ/mol)

q = Heat energy absorbed or released (kJ)

n = number of moles of **solute** or substance used in the reaction (mol).

This can also be written as:

$$\Delta H = -\frac{mc\Delta T}{n}$$

Measuring Molar Heat of Combustion

- Heat liberated when one mole of a substance undergoes complete combustion with oxygen, final products CO2 and water.

The system is the calorimeter, thus to work out the molar heat of combustion:

$$\Delta H = -\frac{q_{solution} + q_{calorimeter}}{n}$$

TOPIC 14 – ENTHALPY AND HESS'S LAW

Law of Conservation of Energy

- Total amount of energy remains constant in any isolated, closed system. Energy can neither be created or destroyed, but changed from one form to another.

Bond Energies

- In a chemical reaction, energy is required to break bonds with the reactant molecules.

Bond breaking – endothermic

Bond making – exothermic

Calculating Enthalpy Change from Average Bond Energies

Example: $H_2 + F_2 \rightarrow 2HF$

Step 1: Draw VSEPR structures for all of the molecules/compounds involved in the reaction.

H – H (Hydrogen gas)

F – F (Fluorine Gas)

H – F (Hydrogen Fluoride)

Step 2: Average Bond Energy Chart: Use the chart to calculate the individual bond energies for the molecules involved.

H – H (432kJ/mol)

F – F (159kJ/mol)

2 H – F (2 x 565kJ/mol = 1130kJ/mol)

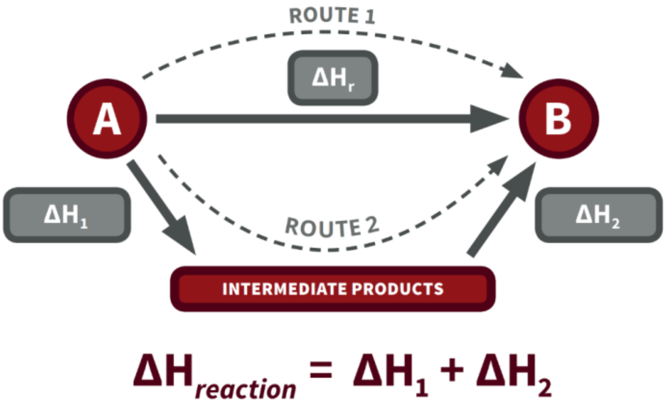
Step 3: Subtract sum of bond energies of products from the reactants.

$$\Delta H = BE_{reactants} - BE_{products}$$

$$\Delta H = (432 + 159) - 1130$$

$$= -539\text{kJ/mol}$$

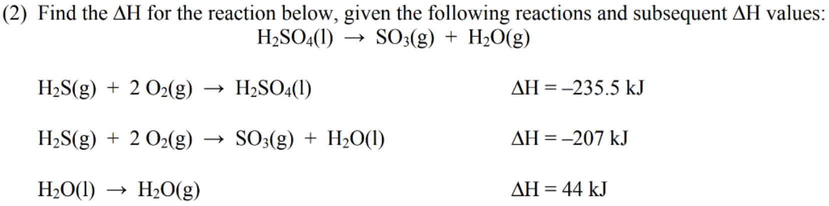
Hess's Law



TIPS FOR CALCULATIONS USING HESS'S LAW:

Question is usually given 1 equation, find enthalpy of that 1 equation given similar equations and their enthalpies.

e.g.



(2) Find the ΔH for the reaction below, given the following reactions and subsequent ΔH values:

$H_2SO_4(l) \rightarrow SO_3(g) + H_2O(g)$

$H_2S(g) + 2\,O_2(g) \rightarrow H_2SO_4(l)$ $\Delta H = -235.5$ kJ

$H_2S(g) + 2\,O_2(g) \rightarrow SO_3(g) + H_2O(l)$ $\Delta H = -207$ kJ

$H_2O(l) \rightarrow H_2O(g)$ $\Delta H = 44$ kJ

TIPS:

- Reactants can only be crossed out with products.
- When reversing a reaction, enthalpy sign switches.
- When multiplying the equation by a particular value, enthalpy is also multiplied by that value.
- Reactants can be subtracted from products and vice versa e.g.

$$3O_2 \rightarrow 5O_2$$

$$\rightarrow 2O_2$$

(taking away $3O_2$ on reactants and minus on products).

- Get the reactants and products needed on their respective side.

Heats of Formation

- Change in enthalpy when one mole of substance in standard state is formed from pure elements under same conditions.

Equation:

$$\Delta H^{\circ}{}_{rxn} = \Sigma \Delta H^{\circ}{}_{f}(prod.) - \Sigma \Delta H^{\circ}{}_{f}(react.)$$

TOPIC 15 – ENTROPY AND GIBBS FREE ENERGY

Entropy

- Measure of a system's thermal energy per unit temperature that is *unavailable* to do work.
- *Measure of how spread out energy is in a system.*

STANDARD ENTROPY CHANGE:

Equation to calculate standard entropy change is:

$$\Delta S^o_{rxn} = \Sigma n S^o(prod.) - \Sigma m S^o(react.)$$

Where:

- S = entropy
- n and m are coefficients found in balanced chemical equation of reaction.

Predicting Entropy Changes

1. Entropy increases with mass

$F_{2(g)}$ = 203 J/K·mol
$Cl_{2(g)}$ = 224 J/K·mol
$Br_{2(g)}$ = 245 J/K·mol
$I_{2(g)}$ = 261 J/K·mol

2. Entropy increases with melting, vaporization or sublimation

$I_{2(s)}$ = 117 J/K·mol vs. $I_{2(\ell)}$ = 261 J/K·mol
$H_2O_{(\ell)}$ = 70 J/K·mol vs. $H_2O_{(g)}$ = 189 J/K·mol

3. Entropy increases when solids or liquids dissolve in water

$CH_3OH_{(\ell)}$ = 127 J/K·mol vs. $CH_3OH_{(aq)}$ = 132 J/K·mol
$NaCl_{(s)}$ = 72 J/K·mol vs. $Na^+_{(aq)} + Cl^-_{(aq)}$ = J/K·mol

4. Entropy decreases when a gas is dissolved in water

$HCl_{(g)}$ = 187 J/K·mol vs. $H^+_{(aq)} + Cl^-_{(aq)}$ = 55 J/K·mol

5. Entropy is lower in hard, brittle substances than in malleable solids like metals

Diamond (C) = 2.4J/mol·K vs. Pb = 65 J/K/mol

6. Entropy increases with chemical complexity

NaCl = 72 J/mol·Kvs. $MgCl_2$ = 90 J/mol·Kvs. $AlCl_3$ = 167 J/K/mol

Predicting Reaction Spontaneity

- Spontaneous process occurs without addition of external energy. Can take quickly or slowly (spontaneity is not related to rate of reaction). Can be *exothermic* or *endothermic*.

Second law of thermodynamics: any spontaneous process must increase entropy in the universe:

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings} > 0$$

Gibbs Free Energy

Equation:

$$\Delta G = \Delta H - T\Delta S$$

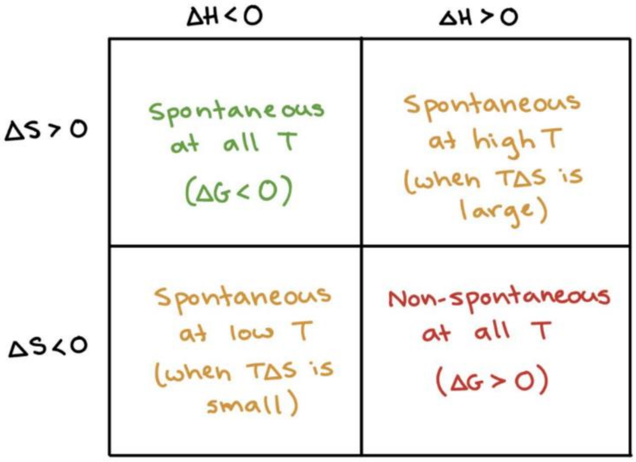
Where:

ΔG = Change in Gibbs free energy (kJ mol^{-1})

ΔH = Change in enthalpy (kJ mol^{-1})

T = temperature (K)

ΔS = Change in entropy (kJ mol^{-1} K^{-1})



	$\Delta H < 0$	$\Delta H > 0$
$\Delta S > 0$	Spontaneous at all T ($\Delta G < 0$)	Spontaneous at high T (when $T\Delta S$ is large)
$\Delta S < 0$	Spontaneous at low T (when $T\Delta S$ is small)	Non-spontaneous at all T ($\Delta G > 0$)

- When $\Delta G < 0$, *exergonic and spontaneous*, produces more products.
- When $\Delta G > 0$, *endergonic and non-spontaneous*, products more starting materials.
- When $\Delta G = 0$, *system in equilibrium*, concentration of reactants/products is the same.